Revision Questions

Module 1

1. What is the difference between primary and secondary research?
2. What does valid mean?
3. What is a hypothesis?
4. What is a survey?
5. What is an experiment?
6. What is a meta study?
7. What is quantitative data?
8. What is qualitative data?
9. What does resolution mean?
10. What does accuracy mean?
11. What is a variable?
12. What does independent variable mean?
13. What does dependent variable mean?
14. What is a control variable?
15. What is an extraneous variable?
16. What are SI units?
17. What points do you need to consider when writing a plan for an investigation?
18. What does “aim” mean in a plan?
19. What is a method?
20. Write a step-by-step method for how to use a separating funnel.
21. Write a step-by-step method for how to carry out an acid-base titration
22. What should you consider when evaluating a method?
23. What goes in the first column of a table?
24. Where does the dependent variable go in a table?
25. What is a data logger?
26. What is a continuous variable? What type of graph should you use to display it?
27. What is a discrete variable? What type of graph should you use to display it?
28. What is a categoric variable? What type of graph should you use to display it?
29. What is standard form?
30. Write 0.000056 in standard form.
31. What are “significant figures”?
32. Write 3467.53 to 3 significant figures
33. An analogue weighing scale is used to measure the mass of an object. The needle is halfway between 0.4g and 0.5g. What figure would you record for the mass?
34. What is a scatter graph?
35. Describe how to draw a line of best fit.
36. How do you calculate the gradient of the slope of a curve?
37. What are concordant results?
38. How do you calculate an average?
39. What is an anomalous result?
40. What should you consider when evaluating results?
41. How can you tell if your results are reliable?
42. What is a false positive?
43. What is a systematic error?
44. What is a random error?
45. What is a margin of error?
46. Describe what “weighing by difference” means.
47. What does precise mean?
48. How is percentage error calculated?
49. I want to measure the temperature rise in an experiment. The starting temperature is 20⁰C and the final temperature is 23⁰C. The margin of error is 0.1⁰C. Calculate the percentage error.

Module 2.1

1. What was believed about atoms at around the 5th century?
2. Describe Dalton’s atomic theory
3. Describe what Joseph John (J.J.) Thomson discovered.
4. Describe the experiment carried out by Ernest Rutherford and explain what it showed.
5. Explain what each of the following scientists contributed to the atomic model – Niels Bohr, Henry Moseley, Rutherford, Louis de Broglie, Erwin Shrodinger and James Chadwick.
6. Describe the current accepted model of the atom.
7. What are the relative masses and charges of protons, neutrons and electrons?
8. Define an isotope
9. Define atomic number and mass number. What are they also known as?
10. What is the overall charge of an atom?
11. Why do different isotopes of the same element react in the same way?
12. What is an ion and how does it form?
13. Define relative isotopic mass.
14. What assumptions do we make about relative isotopic mass?
15. How can we find out the relative isotopic mass?
16. Define relative atomic mass
17. What does the contribution made by an isotope to the overall mass depend on?
18. Calculate the relative atomic mass of Boron which contains 19.77% 10B and 80.23% 11B
19. Define relative molecular mass and explain how we work it out.
20. Define relative formula mass and explain why it is different from relative molecular mass.
21. Give the formula of the following ions:
    1. Sulfate
    2. Ammonium
    3. Hydroxide
    4. Nitrate
    5. Carbonate
    6. Lithium
    7. Aluminium
    8. Chloride
    9. Oxide
22. Define amount of substance
23. Define Avogadro constant.
24. Define the mole.
25. What is the mass of 1 mol of carbon?
26. What is the mass of 10 mol of carbon?
27. What is the mass of 0.5 mol of carbon?
28. Define molar mass, including the units.
29. What is the equation that links amount of substance, molar mass and mass?
30. What is the unit for mass in the above equation?
31. What is the unit for molar mass in the above equation?
32. Define empirical formula
33. Determine the empirical formula of the compound formed when 6.54g of zinc reacts with 1.60g of oxygen.
34. Define molecular formula.
35. Determine the molecular formula of the compound formed from 0.49g of nitrogen combined with 1.12g of oxygen and an Mr of 92.0.
36. Define molar volume
37. At room temperature and pressure, what volume does one mole of gas molecules occupy?
38. What are the units for molar volume?
39. What is the equation that links amount of substance and volume of a gas in dm3? How about in cm3?
40. What amount, in mol, of gas molecules is in 36dm3 of gas?
41. What is the volume of 6 mol of SO2?
42. What is the mass of 0.6 dm3 of N2?
43. What is the volume of 1.282g SO2(g)?
44. What is an “ideal gas”?
45. State the ideal gas equation.
46. What are the units used for pressure, volume and temperature?
47. Calculate the volume of 1 mole of an ideal gas at 0°C and 1 atmosphere pressure.
48. Define concentration and give the units.
49. What is the equation that links amount of substance, volume (in dm3) and concentration? How about when volume is in cm3?
50. What is a standard solution and how do you make one?
51. Find the mass of potassium hydroxide required to prepare 250cm3 of a 0.2 mol dm-3 solution.
52. Define mass concentration. What is the mass concentration of the solution in the previous question?
53. What do the words concentrated and dilute mean?
54. What is a molar solution and what is its symbol?
55. In a chemical reaction, what is conserved?
56. What do the words reactant and product mean?
57. What does the word species mean?
58. What type of species are each of these:
    1. N2
    2. NaCl
    3. C
    4. Na+
    5. He
    6. Na
    7. e-
59. What are the four state symbols and what do they mean?
60. How and why do you balance equations?
61. Balance this equation:

Zn(s) + HNO3(aq) 🡪 Zn(NO3)2(aq) + H2O(l) + NO2(g)

1. Define stoichiometry.
2. What does a balanced equation tell us?
3. A chemist reacted 0.0500 mol of sodium with water to form 50.0 cm3 of aqueous sodium hydroxide.

2Na(s) + 2H2O(l) 🡪 2NaOH(aq) + H2(g)

1. What mass of Na was reacted?

mass = ............................... g

1. Calculate the volume of H2, in dm3, that would be produced at room temperature and pressure, r.t.p.

1 mol of gas molecules occupies 24.0 dm3 at r.t.p.

volume = .......................... dm3

1. Calculate the concentration, in mol dm–3, of NaOH that was formed.

concentration = ................... mol dm–3

1. State the equation for calculating percentage yield
2. What is a yield of 100% rarely reached?
3. 3.925g of 2-chloropropane reacted with an excess of aqueous sodium hydroxide. 2.955g of propan-2-ol was formed as shown in the equation below:

CH3CH(Cl)CH3 + NaOH 🡪 CH3CH(OH)CH3 + NaCl

Calculate the percentage yield of propan-2-ol.

1. Define atom economy
2. State the equation for calculating atom economy
3. What can be done to improve atom economy?
4. 2-iodopropane C3H7I can be hydrolysed using aqueous sodium hydroxide to form the product propan-2-ol. Write a balanced equation for this reaction and calculate the atom economy.
5. Define acid
6. Give the formulae of sulfuric acid, hydrochloric acid and nitric acid.
7. Write an equation to show what happens when HCl(g) is added to an excess of water (use “aq” to show water).
8. Define base
9. Give some examples, including their formula, of different bases.
10. Define alkali and give some examples
11. What is more corrosive – acid or alkali?
12. Write an equation to show what happens when NaOH(s) is added to an excess of water (use “aq” to show water).
13. Write an ionic equation so show what happens when an acid reacts with an alkali. What is the name given to this type of reaction?
14. Describe how ammonia acts as a base and write an equation for the reaction of ammonia with water.
15. Give some examples of biological acids and bases.
16. Define amphoteric and give an example of an amphoteric molecule.
17. Define salt
18. Name the type of salts formed from sulfuric acid, nitric acid and hydrochloric acid.
19. What is special about salts formed from sulfuric acid?
20. Complete these general equations:
    1. Acid + carbonate 🡪
    2. Acid + base 🡪
    3. Acid + alkali 🡪
21. Write balanced symbol equation with state symbols for these reactions:
    1. Hydrochloric acid + calcium carbonate
    2. Sulfuric acid with calcium oxide
    3. Hydrochloric acid with sodium hydroxide
22. What are ammonium salts used for?
23. Write the equation for the reaction of ammonia with nitric acid.
24. Define hydrated and anhydrous.
25. Define water of crystallisation.
26. Write the dot formula for MgCl2H10O5.
27. 6.25g of a hydrated salt, XSO4.5H2O is heated until 4.00g of solid remains. Find the identity of the metal which is a transition element.
28. b) Calculate the percentage of water in the hydrated salt.
29. A hydrate of Sodium Carbonate has a mass of 4.31 g before heating. After heating, the mass of the anhydrous compound is found to be 3.22 g. Determine the formula of the hydrate.
30. What can you find out from a titration?
31. Describe the method for carrying out a titration.
32. Give the names of some indicators that are used and the colour changes that you would observe.
33. 25 cm3 of 0.1 mol/dm-3 sodium hydroxide was exactly neutralised by 20.0cm3 of hydrochloric acid. Calculate the concentration of the acid.
34. A student dissolve 1.571g of an acid, H2X in water and made the solution up to 250cm3. She titrated 25cm3 of this solution again 0.125 mol dm-3 sodium hydroxide, NaOH(aq). 21.30cm3 of NaOH(aq) were needed to reach the end point. The equation for this reaction is:

2NaOH(aq) + H2X(aq) 🡪 Na2X(aq) + 2H2O

Calculate the Mr of the acid H2X.

1. Define oxidation number
2. What is the oxidation number for:
   1. An uncombined element
   2. Combined oxygen
   3. Combined oxygen in peroxides
   4. Combined hydrogen
   5. Combined hydrogen in metal hydrides
   6. A simple ion
   7. Combined fluorine.
3. What is the oxidation number of each atom in NO2?
4. What is the oxidation number of each atom in SO42-?
5. What is the formula for:
   1. Iron(III)oxide
   2. Copper(II)chloride
   3. Nitrate(III)
   4. Sulfate(IV) which has a charge of 2-
   5. Chlorate(V) which has a charge of 1-
6. Define oxidation and reduction in terms of electrons
7. Define redox
8. Write 2 half equations for the reaction of magnesium and chlorine.
9. Define oxidation and reduction in terms of oxidation number
10. Use oxidation numbers to show what has been oxidised and reduced in this reaction:

2Sr(NO3)2(s) → 2SrO(s) + 4NO2(g) + O2(g)

Module 2.2

1. Define shell
2. Define principal quantum number
3. What is the formula for working out the number of electrons from the shell number?
4. Define atomic orbital
5. How many electrons can each orbital hold?
6. What are the four types of orbital?
7. What is the shape of an s-orbital?
8. From n=1 upwards, how many s-orbitals does each shell contain? What is the total number of s electrons in each shell?
9. What is the shape of a p-orbital?
10. From n=2 upwards, how many p-orbitals does each shell contain? What is the total number of p electrons?
11. From n=3 upwards, how many d-orbitals does each shell contain? What is the total number of d electrons?
12. From n=4 upwards, how many f-orbitals does each shell contain? What is the total number of f electrons?
13. How are electrons represented in orbitals?
14. Define sub-shell
15. How are sub-shells filled (the Aufbau principle)
16. Write the electronic configuration of oxygen and draw the electrons in boxes.
17. Write the electronic configuration for zinc.
18. What is special about the 4s and 3d orbitals?
19. How can you use the periodic table to work out the electronic configuration?
20. Explain how noble gases can be used to shorten the electronic configuration.
21. Write the electronic configuration of a Br- ion and an Al3+ ion.
22. Describe the structure of a noble gas atom.
23. Define the octet rule.
24. State three types of chemical bonds.
25. Define compound.
26. Define ionic bond.
27. How does an ionic bond form?
28. Draw dot and cross diagrams to show the bonding in MgCl2
29. Define giant ionic lattice and explain how it forms.
30. How can you use the periodic table to predict an ionic charge?
31. Describe how metals in groups 1-3 form ions
32. Describe how non-metals in groups 15-17 form ions
33. Do Be, B, C and Si usually form ions? Why/why not?
34. Why are oxidation numbers written in roman numbers in the names of some ionic compounds?
35. What is the overall charge of an ionic compound?
36. Predict the formula for lithium nitride, nickel(II) chloride and calcium hydroxide.
37. Define covalent bond.
38. Explain in terms of sub-atomic particles, how the bond forms.
39. Define lone pair.
40. How do single bonds, double bonds and triple bonds form?
41. Draw dot and cross diagrams of CO2, N2, C2H4 and NH3.
42. Define dative covalent/coordinate bonding.
43. Draw a dot and cross diagram of NH4+
44. Draw NH4+using lines and arrows to show the covalent bonds.
45. Draw the dot and cross diagram from the oxonium ion and explain how it forms.
46. Using boron trifluoride as an example, explain what happens if there are no enough electrons to reach an octet.
47. Using sulfur hexafluoride, explain how the octet rule can be expended.
48. What is a better rule than the Octet Rule?
49. Compare the bonding and properties of simple molecular structures and giant covalent compounds.
50. What determines the shape of a molecule or ion?
51. Draw the shapes and bond angles of CH4, BF3, SF6, NH3, H2O, CO2, NH4+ and SO2.
52. Name each of the shapes you have drawn.
53. Define electronegativity.
54. Define permanent dipole
55. Define polar covalent bond
56. Why is CCl4 non-polar?
57. Why is Cl2 non-polar?
58. Why is HCl polar?
59. How is electronegative measured?
60. Complete the sentence: Electronegativity increases \_\_\_\_\_\_\_\_\_\_\_\_\_ a period and \_\_\_\_\_\_\_\_\_\_ a group.
61. If there is a large difference in electronegativity, what type of bond forms?
62. If there is a small difference in electronegativity, what type of bond forms?
63. If the compound is electronically symmetrical, what type of bond forms?
64. Explain why water is polar but carbon dioxide is non-polar.
65. Put these in order of relative strength: ionic bonds, dipole-dipole forces, hydrogen bonds, London dispersion forces, covalent bonds.
66. How are the atoms in a molecule bonded together?
67. Define intermolecular force.
68. Define permanent dipole-dipole force.
69. Draw a diagram to show the dipole-dipole interaction between two HCl molecules.
70. Define London dispersion forces and explain how they form.
71. Why does the boiling point of noble gases increase as you go down the group?
72. Define hydrogen bonding.
73. Draw a diagram to show the hydrogen bonding in water and in ammonia.
74. Describe and explain 2 special properties of water.
75. Describe the importance of hydrogen bonding in biological molecules.

Module 3.1

1. What ideas existed prior to the periodic table?
2. Outline the developments by Antoine-Laurent de Lavoisier, Jons Jakob Berzelius, Johann Wolfgang Dobereiner, John Newlands, Dmitri Mendeleev, Henry Moseley and Glenn Seaborg.
3. Define period and group
4. Define periodicitiy
5. Give some examples of trends that are seen in the periodic table
6. Describe the pattern in electron structure down a group and across a period.
7. Define plasma
8. Define first ionisation energy
9. Write an equation including state symbols for the first ionisation energy of oxygen.
10. Write an equation including state symbols for the second ionisation energy of oxygen.
11. Describe and explain the three factors affecting the size of ionisation energies.
12. Why does each successive ionisation energy get bigger?
13. How can successive ionisation energies give evidence for the presence of shells?
14. Describe and explain the trend in first ionisation energy down a group and across a period.
15. Describe and explain the trend in atomic radii down a group and across a period.
16. Describe, with a diagram, the bonding in metals.
17. Define metallic bonding
18. Define delocalised electrons
19. Define giant metallic lattice.
20. Describe and explain the properties of metals, including their melting/boiling points, electrical conductivity, malleability and ductility.
21. Define, with an example, a metalloid
22. Describe and explain the properties of simple molecular structures.
23. Describe and explain the properties of giant covalent structures.
24. Compare the structure, electrical conductivity and hardness of diamond and graphite.
25. Why is there an increase in boiling point as you move from group 1 to group 3 in period 3?
26. What name is given to group 2 elements?
27. What are the general properties of group 2 elements?
28. Describe the electronic configuration of group 2 elements.
29. Describe and explain the reactivity of group 2 elements.
30. Describe, with an equation, the reaction of a group 2 metal with oxygen. Assign oxidation numbers to your equation and state what is oxidised and what is reduced.
31. Describe, with an equation, the reaction of a group 2 metal with water. Assign oxidation numbers to your equation and state what is oxidised and what is reduced.
32. Why do group 2 oxides and hydroxides react with acids?
33. Give an equation for a group 2 oxide reacting with water.
34. Give an equation to show a solid group 2 hydroxide dissolving in water (use “aq” to symbolise water)
35. Describe the trend in solubility and alkalinity in group 2.
36. Define thermal decomposition
37. Describe the trend in thermal decomposition in group 2.
38. Describe the properties of group 2 elements and their compounds.
39. Describe some uses of group 2 hydroxides.
40. What are group 17 elements called?
41. Describe the physical properties of halogens.
42. Why does the boiling point increase as you go down the group?
43. Describe the electronic configuration of halogens.
44. Explain why reactivity decreases as you move down the group.
45. What colours are produced when chlorine, bromine and iodine molecules are dissolved in water and in cyclohexane?
46. Define displacement reaction.
47. Write ionic equations for the reactions of chlorine with potassium bromide and bromine with sodium iodide.
48. Describe what would be seen in these reactions and explain what is happening.
49. Why won’t iodine react with sodium chloride?
50. Define disproportionation.
51. Write an equation for the reaction of chlorine with water and use oxidation numbers to show this is a disproportionation reaction.
52. Name the two products formed in the above equation.
53. Write an equation for the reaction of chlorine with sodium hydroxide and use oxidation numbers to show this is a disproportionation reaction.
54. Name the two products formed in the above equation.
55. Describe the bonding and structure in halogens.
56. What does diatomic mean?
57. What are the positives and negatives of chlorinating water?
58. Describe the method for testing for halide ions (including the addition of ammonia).
59. What are the positive results for testing for chloride, bromide and iodide ions?
60. Write symbol equations with state symbols for each of the halide tests.
61. Describe the test for carbonate ions
62. Describe the test for sulfate ions
63. Describe the test for ammonium ions
64. Define anion and cation

Module 3.2

1. What is chemical energy?
2. Define enthalpy
3. What do “system” and “surroundings” mean?
4. State the law of conservation of energy.
5. Complete the sentence: “heat loss in a chemical system = heat \_\_\_\_\_\_\_ to the \_\_\_\_\_\_\_\_\_\_\_\_\_\_
6. Complete the sentence: “heat gain in a chemical system = heat \_\_\_\_\_\_\_\_\_ from the \_\_\_\_\_\_\_\_\_\_\_\_\_\_
7. Define exothermic reactions
8. Do exothermic reactions have positive or negative ∆H values?
9. Define endothermic reactions
10. Do endothermic reactions have positive or negative ∆H values?
11. What units are usually stated for energy changes?
12. Draw an energy profile diagram for an exothermic reaction and explain what is happening in terms of energy.
13. Draw an energy profile diagram for an endothermic reaction and explain what is happening in terms of energy.
14. To your profile diagrams, add the activation energy
15. Define activation energy
16. In terms of energy, what does an upwards arrow indicate on an energy profile diagram? What about a downwards arrow?
17. Define thermodynamics
18. State the standard conditions and draw the symbol used to represent this.
19. Define enthalpy change of reaction.
20. Define enthalpy change of formation
21. What is the enthalpy change of formation of an element?
22. Define enthalpy change of combustion
23. Define enthalpy change of neutralisation
24. What is calorimetry?
25. State the expression used to calculate the heat exchanged (q) with the surroundings during calorimetry.
26. Define specific heat capacity
27. Describe how to carry out an experiment to measure the enthalpy change when combusting a fuel.
28. Describe how the carry out an experiment to measure the enthalpy change during a reaction when two solutions are mixed together.
29. Describe a bomb calorimeter.
30. How can you improve a calorimetry experiment?
31. 0.327g of zinc powder was added to 55cm3 of aqueous copper (II) sulfate at 22.8⁰C. The temperature rose to 32.3⁰C. The copper sulfate was in excess. Find the enthalpy change of the reaction, shown by the following equations:

Zn(s) + CuSO4(aq) 🡪 ZnSO4(aq) + Cu(s)

1. Is bond breaking exothermic or endothermic?
2. Is bond making exothermic or endothermic?
3. What is an average bond enthalpy?
4. Calculate the enthalpy of combustion of methane, given that:

C-H = +413

H-O = +464

C=O = +805

O=O = +497

1. Why are average bond enthalpies always positive values?
2. State Hess’ law.
3. What is an enthalpy cycle?
4. State an expression for calculating rate of reaction.
5. State the units usually used in rate of reaction.
6. Describe how the concentration of reactants change during a reaction.
7. Why does rate slow down during a reaction and eventually stop?
8. What factors affect the rate of a reaction?
9. State collision theory.
10. Describe and explain the effects of concentration and pressure of rate.
11. Explain how the rate of reaction can be calculated, including the use of gradients.
12. What can be monitored during a reaction to calculate rate?
13. What is a catalyst and how does it work?
14. What is a homogeneous catalyst?
15. What is a heterogeneous catalyst?
16. Explain the economic importance of catalysts
17. What is the Boltzmann distribution?
18. Draw a sketch of the Boltzmann distribution, remembering to label the axis
19. Explain what the Boltzmann distribution shows.
20. On your sketch, add another curve for a higher temperature and label it
21. To your original curve, show what effect a catalyst has
22. Define reversible reaction, including the symbol used in the equation
23. State the features of a reaction in dynamic equilibrium
24. What does “position of equilibrium” mean?
25. State Le Chatelier’s principle
26. Explain the effect of concentration and pressure on equilibrium
27. Explain the effect of temperature on equilibrium
28. Explain the effect of a catalyst on equilibrium
29. Explain, using the Haber process, the importance of compromise in equilibrium
30. State the conditions in an ammonia plant
31. State the equilibrium law and write a general expression for Kc
32. What do a Kc values of 1, greater than 1 and less than 1 indicate?

Module 4.1

1. Define nomenclature, homologous series, hydrocarbon, aliphatic, alicyclic and aromatic.
2. Name the first 10 unbranched alkanes and state their molecular formula
3. What is the general formula for an alkane and for an alkyl group?
4. Give the formula, prefix (if applicable) and suffix (if applicable) for:
   1. Alcohols
   2. Aldehydes
   3. Alkanes
   4. Alkenes
   5. Carboxylic acids
   6. Haloalkanes
   7. Ketones
5. Draw display and skeletal formula for:
   1. Pent-2-ene
   2. Pentanal
   3. Ethanoic acid
   4. 2-methylhexane
   5. Propanone
   6. Ethan-1,2-diol
   7. Cyclobutane
6. Define unsaturated and saturated
7. Define structural and stereoisomers
8. Explain the Cahn-Ingold-Prelog rules for naming E/Z isomers
9. What is a cis-trans isomer?
10. Draw the unbranched isomers that can be made from C5H10
11. Define heterolytic and homolytic fission and explain what they form.
12. What does a curly arrow show in a mechanism?
13. What is a radical?
14. Describe the bonding in alkanes
15. What are sigma and pi bonds and how do they form?
16. Describe and explain the trend in boiling points of alkanes
17. What affect does branching of alkanes have on boiling point?
18. Write an equation for the complete combustion of methane
19. Suggest an equation for the incomplete combustion of methane
20. Why is incomplete combustion dangerous and what can be done to reduce the risk?
21. What is radical substitution and what are the stages?
22. Why is atom economy often low in radical substitution reactions?
23. Write the equations and label the stages to show the chlorination of ethane to make chloroethane.
24. Describe the bonding in alkenes
25. Why can alkenes form stereoisomers?
26. Define electrophile
27. Why are alkenes more reactive than alkanes?
28. Write an equation for hydrogenation of ethene and state the conditions
29. Write a mechanism for the halogenation of ethene.
30. How can you test if a molecule is saturated?
31. Write an equation for the hydration of ethene and state the conditions
32. Draw a mechanism for the reaction of hydrogen bromide with ethene.
33. Explain the Markownikoff rule.
34. Draw the major product of the reaction between propene and hydrogen chloride.
35. Define monomer and polymer
36. What is a repeating unit?
37. Draw the repeating unit for tetrafluorethene
38. Draw the monomer used to make poly(propene)
39. State the different ways of dealing with polymer waste and the issues associated with each one.
40. Define biodegradable, photodegradable and bioplastic

Module 4.2

1. What are primary, secondary and tertiary alcohols?
2. Describe and explain the trend in boiling points of alcohols
3. Describe and explain the volatility of alcohols compared to alkanes
4. Describe and explain the solubility of alcohols in water.
5. Write an equation for the complete combustion of propanol in oxygen.
6. Describe the oxidation of primary, secondary and tertiary alcohols. State the conditions and the oxidising agent.
7. Write equations for the partial oxidation of propan-1-ol using [O] to represent the oxidising agent.
8. Write equations for the complete oxidation of propan-1-ol using [O] to represent the oxidising agent.
9. Write equations for the oxidation of propan-2-ol using [O] to represent the oxidising agent.
10. What is observed during oxidation?
11. What apparatus is used?
12. What is an ester and how is it formed?
13. Name and draw the ester formed from the reaction between butan-1-ol and propanoic acid.
14. Define dehydration
15. Define elimination reaction
16. Write an equation for the dehydration of ethanol
17. Give the conditions needed for dehydration of alcohols
18. How do you carry out halide substitution of an alcohol using chloride ions?
19. How do you carry out halide substitution of an alcohol using bromide ions?
20. How do you carry out halide substitution of an alcohol using iodide ions?
21. What are haloalkanes?
22. Define nucleophile
23. Why are haloalkanes polar?
24. What happens to the strength of a C-X bond (where X is a halide) going down the halogen group and why?
25. Define nucleophilic substitution
26. Define hydrolysis
27. Draw the mechanism for the hydrolysis of 1-iodopropane
28. Outline an experiment that could be used to measure the rate of hydrolysis
29. Explain why the rate of hydrolysis is fastest in iodo-alkanes than fluor-alkanes
30. What is a chlorofluorocarbon?
31. What were CFCs used for?
32. What problem do CFCs cause?
33. Write the steps involved in radical substitution of ozone to break down O3 to O2.
34. Define ozone layer
35. Write the steps involved in radical substitution of ozone to break down O3 to O2 using nitrogen oxides.
36. What is the Montreal Protocol?
37. What is Quickfit apparatus?
38. State the apparatus needed for distillation and explain how to set it up and when to use it.
39. State the apparatus needed for reflux and explain how to set it up and when to use it.
40. Describe how to prepare and purify an organic liquid.
41. Explain how to test for:
    1. Unsaturated hydrocarbons
    2. Haloalkanes
    3. Carbonyl groups
    4. Carboxylic acids
    5. Alcohols
42. What effects the frequency of vibration when molecules absorb infrared radiation?
43. What can infrared spectroscopy be used for?
44. What are the units for frequency in infrared spectroscopy?
45. You have 3 samples – one is propan-1-ol, one is propanone and other other is propanoic acid. How could infrared spectroscopy be used to positively identify each one?
46. How does a mass spectrometer work?
47. Define molecular ion and m/z.
48. Suggest the identity of peaks at:
    1. 15
    2. 17
    3. 29
    4. 43
    5. 57
49. What is produced when the molecular ion fragments?
50. Why is there usually a small peak to the right of the molecular ion peak?
51. An analytical chemist was provided with a compound **J** which has an unbranched carbon skeleton. After analysis, the chemist obtained the following results:

|  |  |
| --- | --- |
| **type of analysis** | **evidence** |
| infrared spectroscopy | broad absorption at 3350 cm–1 |
| percentage composition by mass | C, 70.59%; H, 13.72%; O, 15.69% |
| mass spectrometry | molecular ion peak at *m*/*z* = 102.0 |

Use this information to suggest all the possible structures for the **unbranched** compound **J.**